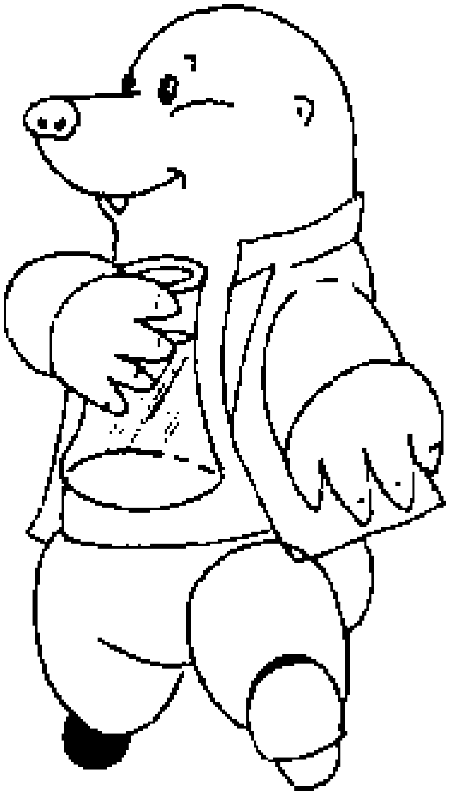
Chemical Bonding

Downingtown East High School
Science Department

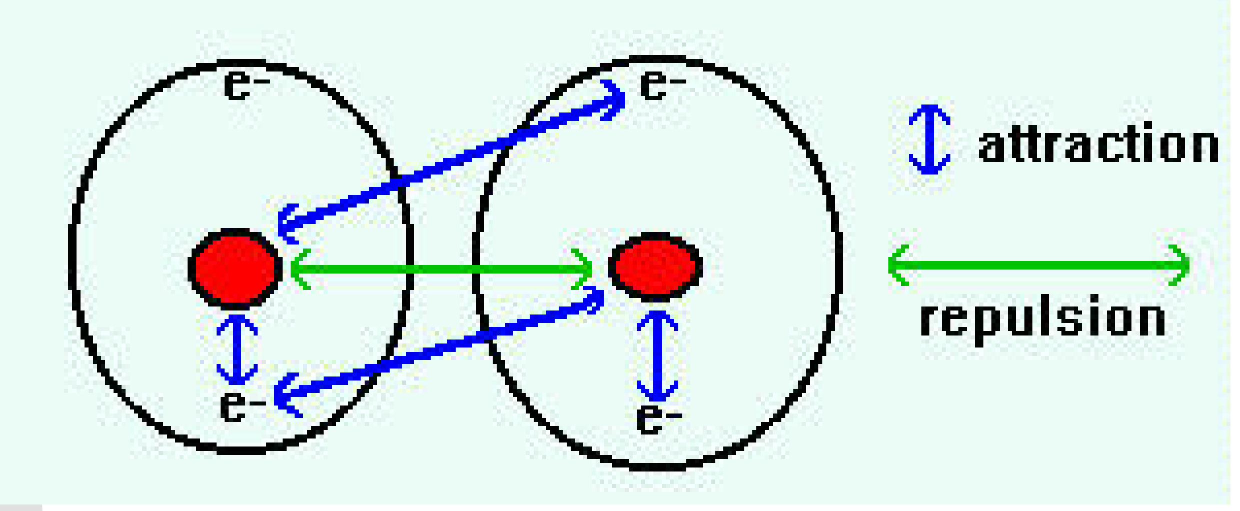


Chemical Bonding

- All bonding forces are due to **electrostatic charge**.
- Opposite charges attract, Like charges repel.

Chemical Bonding

- This diagram shows the attraction and repulsion between atoms:The outer ring (e-) is the *electron cloud*. The inner red ring is the *nucleus*



Metallic Bonding

- Bonding between atoms with *low electronegativity.*ie 1,2 or 3 valence electrons, therefore there are many *vacancies in valence shell.*
  When electron clouds overlap, *electrons can move into electron cloud of adjoining atoms.*

Metallic Bonding

- Each atom becomes surrounded by a number of others in a 3-D lattice, where valence electrons move freely from 1 valence shell to another.

- **Delocalised** valence electrons moving between nuclei *generate a binding force to hold the atoms together*

Metallic Bonding

- This table shows metallic properties and the explanations of these properties.

Observation	***Explanation***
Metals are dense	The particles present in metals are tightly packed in the lattice.
Metals have high melting and boiling points.	Strong forces of attraction exist between particles. A Large amount of thermal energy is required to overcome the strong electrical forces between the positive ions and the delocalised electrons. These forces operate throughout the lattice.
Metals are good conductors of heat.	Delocalised electrons transmit the energy of vibrations of 1 positive ion to its neighbors.
Metals are good conductors of electricity.	Mobile delocalised electrons within the lattice. Electrons flow in at one end, and the same number flow out the other end.
Metals are malleable and ductile.	distortion does not disrupt the metallic bonding.
Metals are lustrous.	presence of free electrons causes most metals to reflect light (non-metals are transparent).

Ionic Bonding

- occurs between *metals* and *non-metals*..
  - **Metal** atoms have a *low number of valence electrons* and a *low electronegativity.*
  - **Non-metal** atoms have *numerous valence electrons*.

Ionic Bonding

If the electron clouds overlap (bond)..

Metals
lose valence electrons
achieve a stable valence shell (usually 8 e-)
gains a positive charge, ie a positive ion.
Non-metals
gain valence electrons
achieve a stable valence shell (usually 8 e-)

Ionic Bonding

eg:

- Na atom [Ne] $3s^1$ goes to the sodium ion(Na^{+1}) with [He] $2s^2 2p^6$
  K atom [Ar] $4s^1$ goes to the potassium ion (K^{+1}) with [Ne] $3s^2 3p^6$
  Mg atom [Ne]$3s^2$ goes to the magnesium ion (Mg^{+2}) with [He] $2s^2 2p^6$
  Ca atom [Ar] $4s^2$ goes to the calcium ion (Ca^{+2}) with [Ne] $3s^2 3p^6$
- Al atom [Ne] $3s^2 3p^1$ goes to the aluminum ion (Al^{+3}) with [He] $2s^2 2p^6$
- The charge is the number of valence e- it has to lose.

Ionic Bonding

eg:

- F atom [He] $2s^2 2p^5$ goes to the fluoride ion (F^{-1}) with [He] $2s^2 2p^6$
- Cl atom [Ne]$3s^2 3p^5$ goes to the chloride ion (Cl^{-1}) with [Ne] $3s^2 3p^6$
- O atom [He] $2s^2 2p^4$ goes to the oxide ion (O^{-2}) with [He] $2s^2 2p^6$
- The charge is the number of electrons needed to get 8 in the outermost level and make the number negative.
  - Ex: O gains 2 e^- and has –2 charge

Ionic Lattice

- *Positive and negative ions attract each other to form a three dimensional continuous lattice* structure.
  - Each positive ion is surrounded by a number of negatives.
  - Each negative ion is surrounded by a number of positives.
  - The ratio of positive to negative ions in the lattice is determined by the charges of the ions.

Ionic Lattice

- This table shows properties of ionic lattices (compounds) and explanations of these properties

Property	Explanation
Melting point and boiling point	The melting and boiling points of ionic compounds are high because a large amount of thermal energy is required to seperate the ions which are bound by strong electrical forces.
Electrical conductivity	Solid ionic compounds do not conduct electricity when a potential is applied because there are no mobile charged particles. No free electrons causes the ions to be firmly bound and cannot carry charge by moving.
Hardness	Most ionic compounds are hard; the surfaces of their crystals are not easily scratches. This is because the ions are bound strongly to the lattice and aren't easily displaced.
Brittleness	Most ionic compounds are brittle; a crystal will shatter if we try to distort it. This happens because distortion cause ions of like charges to come close together then sharply repel.

Ionic Lattice

- The table below shows examples of some ionic compounds.

Formula	**Name**
Na_2CO_3	Sodium Carbonate
KI	Potassium Iodide
$AlCl_3$	Aluminium Chloride
$Ca_3(PO_4)_2$	Calcium Phosphate
$Zn(CH_3COO)_2$	Zinc Acetate
$Mg(OH)_2$	Magnesium Hydroxide
$(NH_4)_2SO_4$	Ammonium Sulphate
$AgNO_3$	Silver Nitrate
$KHCO_3$	Potassium Hydrogencarbonate
PbS	Lead (II) Sulphide
$SnCl_4$	Tin (IV) Chloride

Covalent Bonding

- Bonding between *non-metals* and *non-metals*.
  - Therefore all atoms included have *fairly high electronegativity* and *few vacancies in valence energy levels*.
  - When they bond, they gain electrons to achieve stable configuration.
  - Hence, electrons are **shared**.

Covalent Bonding

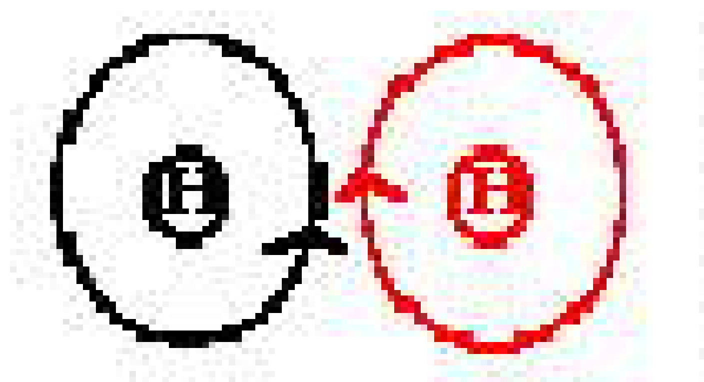
- Sharing produces low energy (stable) electron arrangements.
- *ie:*
  - Full outer shell (eg He $1s^2$; Ne $1s^2\ 2s^2\ 2p^6$) 8 electrons (4 pairs) in outer shell (eg Ar $1s^2\ 2s^2\ 2p^6\ 3s^2 3p^6$)

The octet rule

- The **octet rule** says that atoms tend to gain, lose or share electrons so as to have eight electrons in their outer electron shell.
- It is a very useful rule but you should also know that there are *many* bonding situations where it does *not* apply.
- As you learn to use the octet rule, also learn to recognize situations where it does not apply and disregard it in those situations.

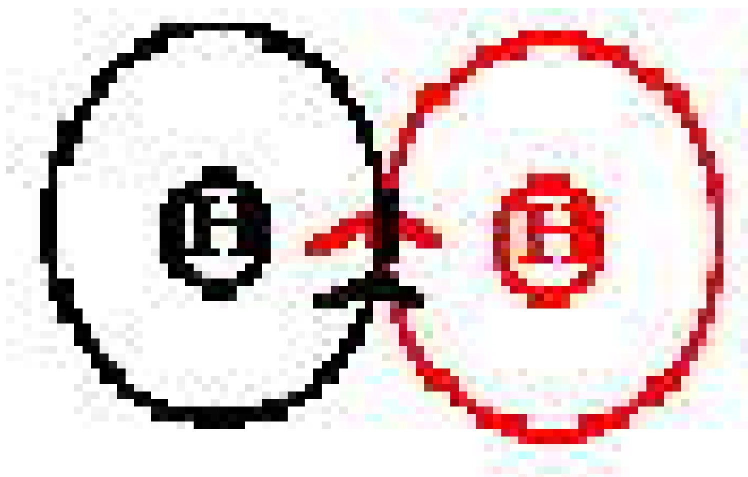
Covalent Bonding

- **Covalence** is the number of electrons an atom needs to gain to produce a stable outer shell. The number of *shared pairs (covalent bonds)* of electrons an atom forms. eg Hydrogen *H* needs 1 additional electron



Covalent Bonding

- Hydrogen bonds1 pair of e-'s shared between 2 atoms - *covalent bond*.
  -

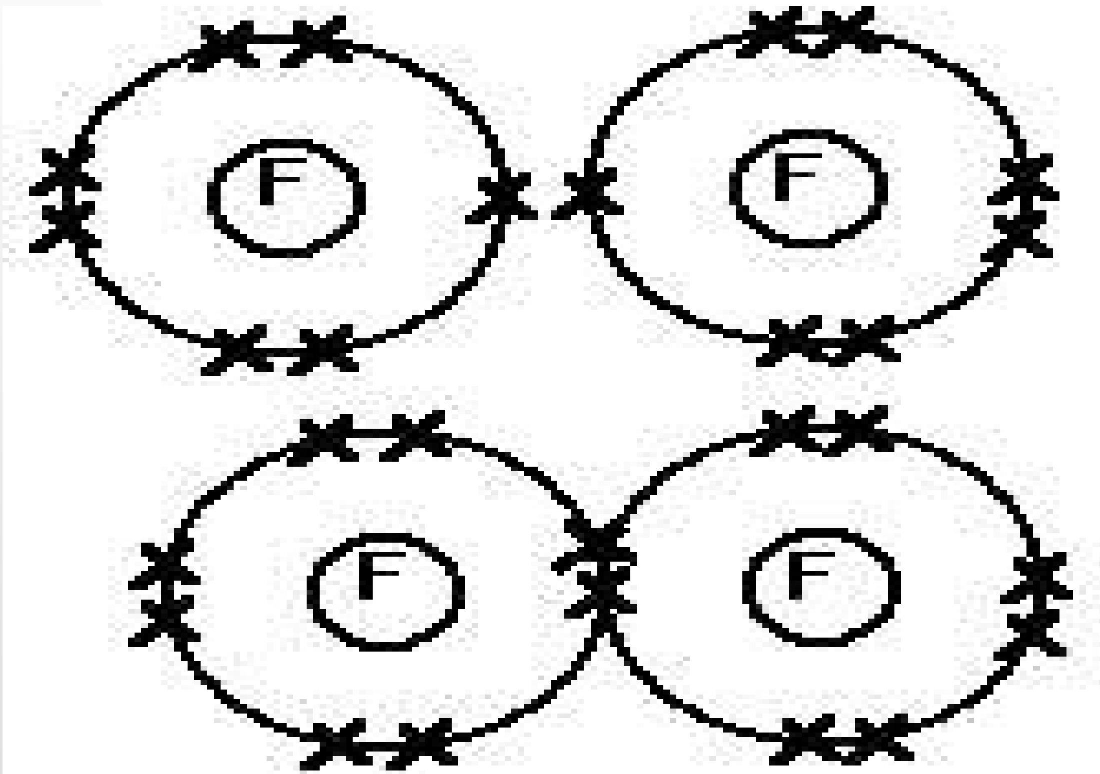


Covalent Bonding

- Bonding pairs of electrons orbits both nuclei - attracts both nuclei - provides bonding force.
- Hydrogen molecule consists of 2 covalently bonded hydrogen atoms, which have no tendency to bond further (both have achieved a stable outer shell).
- Each molecule exists independently.

Covalent Bonding

- eg Fluorine F [He] $2s^2 2p^5$ *Covalence = 1*



Covalent Bonding

- This table shows properties of covalent molecular compounds and explanations of these properties

Property	**Explanation**
Do not conduct electricity.	No mobile charged particles Molecules not charged Electrons tightly bound to atoms or shared by atoms in covalent bonds.
Melting and boiling points low.	During melting/boiling, molecules become separated. Forces of attraction between molecules are weak and little thermal energy is required to separate them.
Soft.	Molecules weakly attracted to each other and are easily displaced.

Covalent Bonding

- In covalent bonds, electron pairs are shared equally between atoms of equal electronegativity.
- If the atoms in a covalent bond have differing electronegativities, the atoms with the higher electronegativity has >50% of the shared pairs of electrons and the atoms with low electronegativity has 50% of the shared pairs of electronegativity.

Covalent Bonding

- The atom tending to gain electrons acquires a slight *negative charge* (delta negative)
- The atom tending to lose electrons acquires a slight *positive charge* (delta positive)
- *The bond is* **POLAR.**

Expansion of the Octet

- Atoms, which have room for more than 8 electrons in their outer shell, may form covalent bonds, which result in 10, 12, or 14 outer shell electrons (even numbers are stable).

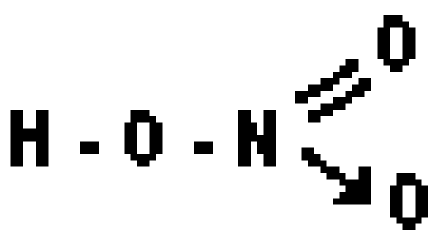
Expansion of the Octet

- ie atoms in Period 3 or below are able to form shared pairs -> even number of e- results
  - (room for 18 valence e-)
- eg:
  - P [Ne]$3s^2 3p^3$ - 3 shared pairs, 8 valence e-'s, or 5 shared pairs, 10 valence e-'s.

  - S [Ne]$3s^2 3p^4$ - 2 shared pairs, 8 valence e-'s, or 4 shared pairs, 10 valence e-'s, or 6 s.p, 10 v.e-'s.

Co-ordinate Bonds

- A covalent bond where the shared pair of e-'s is provided by 1 atom.
  It is represented by an arrow, showing the direction in which the shared e-'s are provided.
  - eg HNO_3

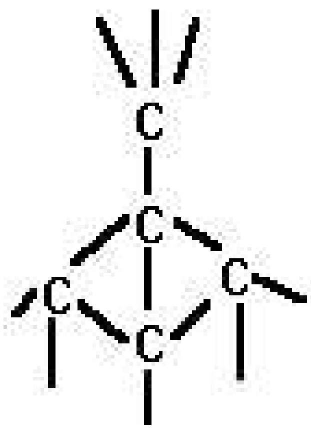
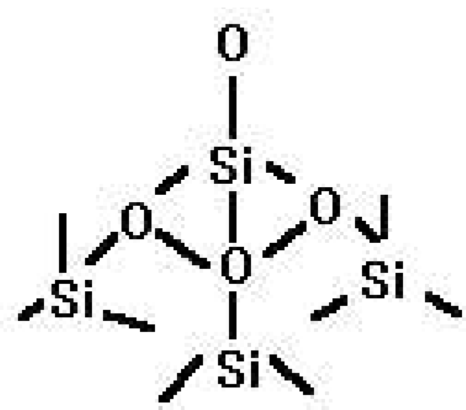


  - Co-ordinate bonds aren't very common. They usually occur in atoms that cannot expand the octet.

Covalent Network (Lattice)

- Bonding between non-metals.
  Form covalent bonds (shared electron pairs).
- Do not form separate (discrete) molecules but a *continuous network*.

Covalent Network (Lattice)

- eg Diamond (carbon) ->   - Silicon Dioxide SiO_2.

Covalent Network (Lattice)

- *This table shows properties of covalent network substances and explanations of these properties.*

Property	**Explanation**
Electrical conductivity	Poor conductors because electrons are held either on the atoms or within covalent bonds. They cannot move through the lattice. Graphite is an exception.
Melting points	Very high melting points because each atom is bound by strong covalent bonds. Many covalent bonds must be broken if the solid is to be melted and a large amount of thermal energy is required for this.
Hardness	They are hard because the atoms are strongly bound in the lattice, and are not easily displaced.
Brittleness	Covalent network substances are brittle. If sufficient force is applied to a crystal, covalent bond are broken as the lattice is distorted. Shattering occurs rather than deformation of a shape.

Resonance

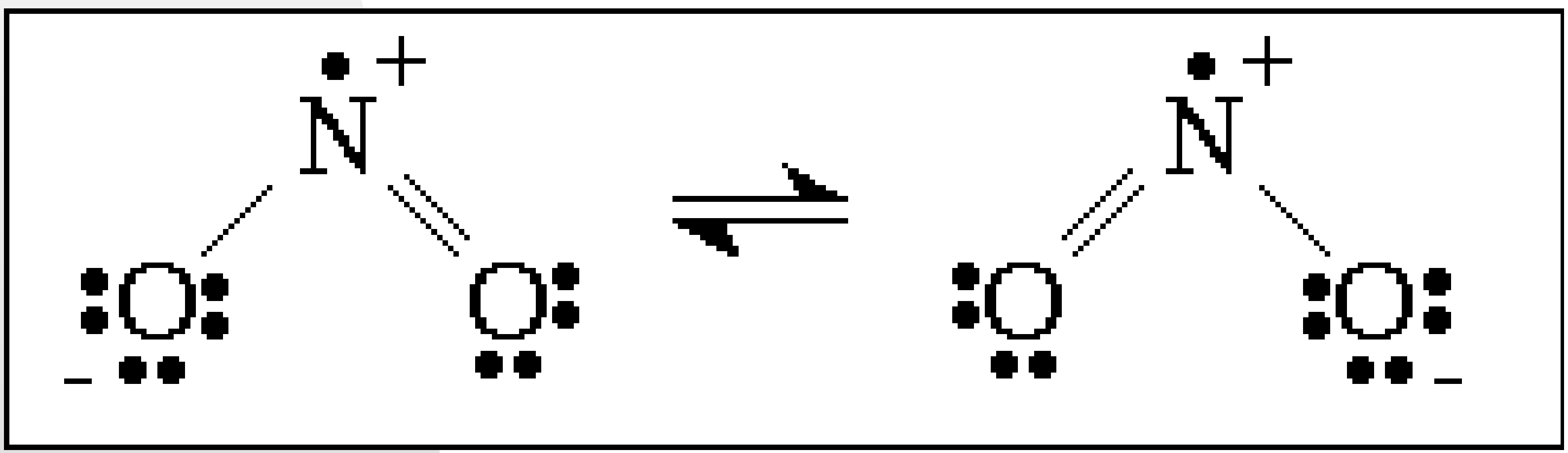
- There are a number of compounds and polyatomic ions that cannot be written using one single structure.
- Linus Pauling developed what today is called "resonance theory."

Resonance

- **Resonance happens when more than one valid Lewis dot-diagram (or what Pauling calls a valence-bond structure) can be written for a molecule or ion. When this happens, the true structure is a blend of all the different possible structures.**

Resonance

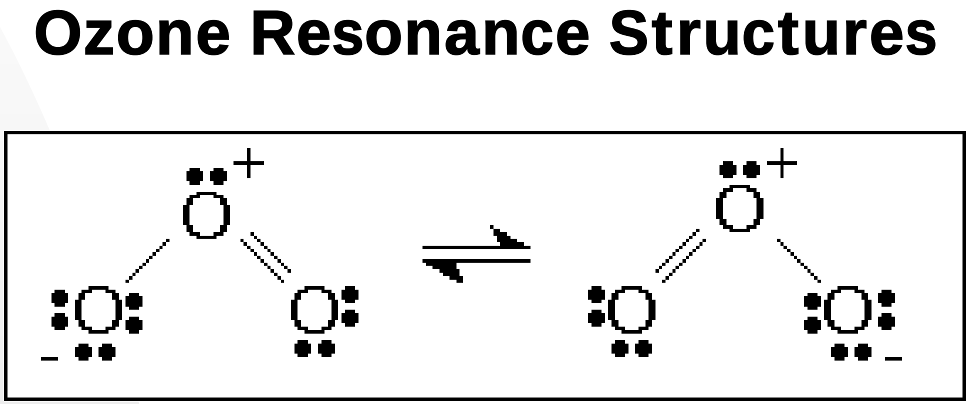
- Here is another example, using the molecule NO_2:



Resonance

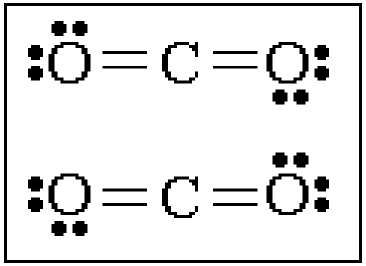
- OK, here's the deal. Neither one of those two structures really does exist.
- The real molecule that exists in nature is a "resonance hybrid" between the two.
- The real molecule acts as if it had one and one-half bonds between each of the two structures.
- The two structures above are merely descriptive aids and, in fact, never exist.

Resonance



Resonance

Carbon Dioxide Resonance Structures



Intermolecular Forces

- Ionic Compounds and Metals
- **Electrostatic forces -** these forces occur between charged species and are responsible for the extremely high melting and boiling points of ionic compounds and metals.

Intermolecular Forces

- Covalent Compounds
- **London forces -** all molecules have the capability to form London forces. These are solely dependent on the surface area and the polarizability of the surface of the molecule. These are the only types of forces that non-polar covalent molecules can form. They result from the movement of the electrons in the molecule, which generates temporary positive and negative regions in the molecule.

Intermolecular Forces

- **Dipole-dipole forces -** only polar covalent molecules have the ability to form dipole-dipole attractions between molecules. Polar covalent molecules act as little magnets; they have positive ends and negative ends that attract each other.
- **Hydrogen bonding -** these occur between polar covalent molecules that possess a hydrogen bonded to an extremely electronegative element, specifically - N, O, and F. Water is a major example of this type of bonding and explains why the mp and bp of water are so high.